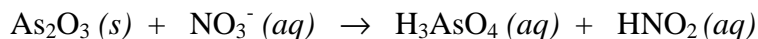


## Half-Reaction Method for Balancing Redox Reactions

This method is based on separating a redox reaction into an oxidation process and a reduction process. Each process (half-reaction) is balanced independently and then the half-reactions are combined to obtain the final balanced overall redox reaction, the net ionic equation.

Consider the reaction between arsenic oxide and nitrate ion in acidic solution:



Assign oxidation numbers to all reactants and products until you find the species that exchange electrons.

1. Identify which species gets oxidized and which gets reduced then write the half-reactions:

Arsenic (as  $\text{As}_2\text{O}_3$ ) loses 2 electrons to give arsenic acid ( $\text{H}_3\text{AsO}_4$ ):



(phase labels will be omitted for clarity)

Nitrogen (as nitrate ion) gains 2 electrons to give nitrous acid:



notice that oxygen does not undergo redox in this reaction

2. Balance each half-reaction for all atoms; balance oxygen and hydrogen last.

Balance arsenic and add 2 electrons per atom of arsenic, then add  $\text{H}_2\text{O}$  to balance oxygen and finally add  $\text{H}^+$  to balance hydrogen:

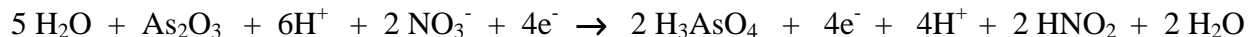
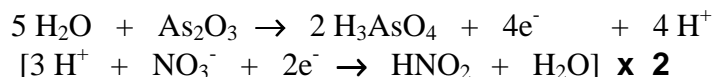


Nitrogen is already balanced so add  $\text{H}_2\text{O}$  then  $\text{H}^+$  to balance oxygen and hydrogen:

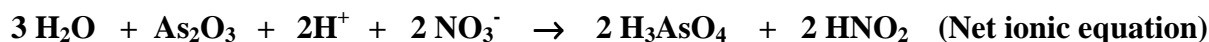


3. Multiply each half-reaction by some factor so that the number of electrons lost in the oxidation is equal to the number of electrons gained in the reduction. Then add together the half-reactions and cancel all species common to both sides of the overall equation:

The reduction half-reaction must be multiplied by **two** to reconcile the electron exchange:



These species,  $2 \text{H}_2\text{O}$ ,  $4 e^-$ ,  $4 \text{H}^+$ , cancel from both sides to give the net ionic equation:



(3  $\text{H}_2\text{O}$  remain on the reactant side of the net ionic equation)

*For redox reactions that occur in basic solution, there is one additional step to the half-reaction method: See the next example...*

